Name: KEY Off. Class: _____ Per: _____ Date: _____

Teacher: _____ **Heat Calculation Regents Practice** **Chemistry**

A 1. As a 15.1-gram sample of a metal absorbs 48.75 J of heat, its temperature increases 25.0K. What is the specific heat capacity of the metal? $q = mc\Delta T$

A) 0.129 J/g°K B) 1.95 J/g°K
C) 3.23 J/g°K D) 7.74 J/g°K

B 2. An 80.0-gram sample of water at 10.0°C absorbs 1680 Joules of heat energy. What is the final temperature of the water? $q = mc\Delta T$

A) 50.0°C B) 15.0°C
C) 5.00°C D) 4.00°C

B 3. When 200 grams of water cools from 50.°C to 25°C, the total amount of heat energy released by the water is $q = mc\Delta T$

A) 42 kJ B) 21 kJ C) 34 J D) 17 J

C 4. How much heat energy must be absorbed to completely melt 35.0 grams of $H_2O(s)$ at 0°C? $q = mH_f$

A) 9.54 J B) 146 J
C) 11 700 J D) 79 100 J

B 5. At 1 atmosphere of pressure, 25.0 grams of a compound at its normal boiling point is converted to a gas by the addition of 34,400 Joules. What is the heat of vaporization for this compound, in Joules per gram? $q = mH_v$

A) 25.0 J/g B) 1376 J/g
C) 2,260 J/g D) 34,400 J/g

C 6. What is the amount of heat required to completely melt a 200.-gram sample of $H_2O(s)$ at STP? $q = mH_f$

A) 334 J B) 836 J
C) 66800 J D) 452000 J

B 7. The temperature of a sample of water changes from 10°C to 20°C when the sample absorbs 418 joules of heat. What is the mass of the sample? $q = mc\Delta T$

A) 1 g B) 10 g
C) 100 g D) 1000 g

C 8. What is the total amount of heat required to completely melt 347 grams of ice at its melting point? $q = mH_f$

A) 334 J B) 1,450 J
C) 116,000 J D) 784,000 J

D 9. The heat of vaporization of a liquid is 1,340 Joules per gram. What is the minimum number of Joules needed to change 40.0 grams of the liquid to vapor at the boiling point? $q = mH_v$

A) 33.5 B) 1,340
C) 3,280 D) 53,600

C 10. A 36-gram sample of water has an initial temperature of 22°C. After the sample absorbs 1200 joules of heat energy, the final temperature of the sample is $q = mc\Delta T$

A) 8.0°C B) 14°C
C) 30.°C D) 55°C

C 11. The temperature of a sample of water changes from 10.°C to 20.°C when the water absorbs 420 Joules of heat. What is the mass of the sample? $q = mc\Delta T$

A) 1.0 g B) 10. g
C) 100 g D) 1000 g

B 12. How many joules of heat are absorbed to raise the temperature of 435 grams of water at 1 atm from 25°C to its boiling point, 100.°C? $q = mc\Delta T$

A) 4.5×10^4 J B) 1.4×10^5 J
C) 2.5×10^7 J D) 7.4×10^7 J

B 13. What is the total number of joules released when a 5.00-gram sample of water changes from liquid to solid at 0°C? $q = mH_f$

A) 334 J B) 1670 J
C) 2260 J D) 11 300 J

C 14. How many Joules of heat energy are released when 50. grams of water are cooled from 70.°C to 60.°C? $q = mc\Delta T$

A) 42 J B) 210 J
C) 2100 J D) 4200 J

D 15. How much energy is required to vaporize 10.00 grams of water at its boiling point? $q = mH_v$

A) 2.26 kJ B) 3.34 kJ
C) 4.2 kJ D) 22.6 kJ

. short answer questions on back

see attached for work for #1-15

KEY

16. Base your answer to the following question on the information below and on your knowledge of chemistry.

Fruit growers in Florida protect oranges when the temperature is near freezing by spraying water on them. It is the freezing of the water that protects the oranges from frost damage. When $H_2O(\ell)$ at 0°C changes to H_2O (s) at 0°C, heat energy is released. This energy helps to prevent the temperature inside the orange from dropping below freezing, which could damage the fruit. After harvesting, oranges can be exposed to ethene gas, C_2H_4, to improve their color.

Determine the quantity of heat released when 2.00 grams of H_2O (l) freezes at 0°C.

$q = mH_f$ $q =$ $m = 2.00g$ $H_f = 334\ J/g$ (Table B)

$q = (2.00g)(334\ J/g)$

$q = 668\ J$

17. Show a numerical setup for calculating the quantity of heat in joules required to completely vaporize 102.3 grams of $H_2O(\ell)$ at 100.°C and 1.0 atm.

$q = mH_v$ $q = ?$ $m = 102.3\ g$ $H_v = 2260\ J/g$ (Table B)

$q = (102.3\ g)(2260\ J/g)$

set-up only!

Base your answers to questions **18** and **19** on the information below.

Heat is added to a 200.-gram sample of $H_2O(s)$ to melt the sample at 0°C. Then the resulting $H_2O(\ell)$ is heated to a final temperature of 65°C.

18. Compare the amount of heat required to vaporize a 200.-gram sample of $H_2O(\ell)$ *at its boiling point to the amount of heat required to melt a 200.-gram sample of $H_2O(s)$ at its melting point.*

- more heat is required to vaporize water
- less heat is needed to melt water

Think! H_f is lower than H_v

19. In the space below, show a numerical setup for calculating the total amount of heat required to raise the temperature of the $H_2O(\ell)$ from 0°C to its final temperature. (65°C)

$q = mC\Delta T$ $q = ?$ $m = 200.g$ $C = 4.18\ J/g°C$ $\Delta T = 65°C$ $T_f = 65°C$ $T_i = 0°C$

$q = (200g)(4.18\ J/g°C)(65°C)$

set-up only

20. Base your answer to the following question on the information below and you knowledge of chemistry.

A 100.-gram sample of liquid water is heated from 20.0°C to 50.0°C. Enough $KClO_3(s)$ is dissolved in the sample of water at 50.0°C to form a saturated solution.

Using the information on Table *B*, determine the amount of heat absorbed by the water when the water is heated from 20.0°C to 50.0°C.

$q = mC\Delta T$

$q =$ $m = 100g$ $C = 4.18\ J/gC$ (TABLE B) $\Delta T = 50 - 20 = 30°C$ $T_f = 50.0°C$ $T_i = 20.0°C$

$q = (100g)(4.18\ J/gC)(30°C)$

$q = 12540\ J$

Heat Calculation Regents Practice - KEY

1. $q = mC\Delta T$ (Unit of temperature doesn't matter)

$q = 48.75\,J$

$m = 15.1\,g$

$C = ?$

$\Delta T = 25.0\,K$

$T_f = -$

$T_i = -$

$48.75\,J = (15.1\,g)(c)(25.0\,K)$

$\frac{48.75\,J}{377.5\,gK} = \frac{377.5\,gK\,(c)}{377.5\,gK}$

$\boxed{0.129\,\frac{J}{gK} = c}$ (A)

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2. $q = mC\Delta T$

$q = 1680\,J$

$m = 80.0\,g$

$C = 4.18\,J/g^\circ C$ (TABLE B)

$\Delta T =$

$T_f = ?$

$T_i = 10.0^\circ C$

$1680\,J = (80.0\,g)(4.18\,\frac{J}{g^\circ C})(\Delta T)$

$\frac{1680\,J}{334\,J/C} = \frac{334\,J}{334\,C}$

$5.029 = \Delta T$

$\Delta T = T_f - T_i$

$5.03 = T_f - 10.0^\circ C$

$+10.0 \qquad +10.0^\circ C$

$\boxed{15.03 = T_f}$ (B)

3. $q = mc\Delta T$

$q = ?$

$m = 200.\ g$

$c = 4.18\ J/g°C$ (TABLE B)

$\Delta T = 25°C - 50°C = -25°C$

$T_f = 25°C$

$T_i = 50°C$

"-" temperatures means it was cooled - this OK!

$q = (200.\ g)(4.18\ \frac{J}{g°C})(-25°C)$

$q = -20900\ J$

★ answer choices in KJ = $-20.9\ KJ$

choice (B) 21 KJ

(Remember "-" means heat is released - answer choices won't have a negative choice b/c the question says "released"

4. $q = mH_f$

$q = ?$

$m = 35.0\ g$

$H_f = 334\ J/g$ (TABLE B)

$q = (35.0\ g)(334\ \frac{J}{g})$

$q = 11690\ J \rightarrow 11700\ J$ (C)

5. $q = mH_v$

$q = 34{,}400\ J$

$m = 25.0\ g$

$H_v = ?$

$\frac{34{,}400\ J}{25.0\ g} = \frac{25.0\ g \cdot H_v}{25.0\ g}$

$1376\ J/g = H_v$ (B)

6. $q = mH_f$
$q = ?$
$m = 200\,g.$
$H_f = 334\,J/g$ (TABLE B)

$q = (200\,g)(334\,J/g)$
$\boxed{q = 66800\,J}$ (C)

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7. $q = mc\Delta T$
$q = 418\,J$
$m = ?$
$c = 4.18\,J/g°C$ (TABLE B)
$\Delta T = 20 - 10 = 10°C$
$T_f = 20°C$
$T_i = 10°C$

$418\,J = m\left(4.18\,\frac{J}{g}\right)(10°C)$

$\frac{418}{41.8} = \frac{m\,41.8}{41.8}$

$\boxed{10.\,g = m}$ (B)

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8. $q = mH_f$
$q = ?$
$m = 347\,g$
$H_f = 334\,J/g$

$q = (347\,g)\left(334\,\frac{J}{g}\right)$

$q = 115898\,J$
$\hookrightarrow \boxed{116000\,J}$ (C)

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9. $q = mH_v$
$q = ?$
$m = 40.0\,g$
$H_v = 1340\,J/g$

$q = (40.0\,g)(1340\,J/g)$
$\boxed{q = 53600\,J}$ (D)

→

10. $q = mc\Delta T$

$q = 1200\,J$

$m = 36.\,g$

$c = 4.18\,J/g°C$ (TABLE B)

$\Delta T =$

$T_f = ?$

$T_i = 22°C$

$1200 = (36)(4.18)(\Delta T)$

$\frac{1200}{150.48} = \frac{150.48}{150.48}\,\Delta T$

$7.97 = \Delta T$

$\hookrightarrow 8°C = \Delta T$

$\Delta T = T_f - T_i$

$8°C = T_f - 22°C$

$+22 \qquad +22$

$\boxed{T_f = 30°C}$ (C)

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11. $q = mc\Delta T$

$q = 420\,J$

$m = ?$

$c = 4.18\,J/g°C$ (TABLE B)

$\Delta T = 20°C - 10°C = 10°C$

$T_f = 20°C$

$T_i = 10°C$

$420\,J = m\,(4.18\,J/g)(10°C)$

$\frac{420\,J}{4.18} = m\,\frac{41.8\,J°C}{g}$

$\frac{4.18}{4.18}$

$\boxed{m = 100.47\,g}$ (C)

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12. $q = mc\Delta T$

$q = ?$

$m = 435\,g$

$c = 4.18\,J/g°C$ (TABLE B)

$\Delta T = 100°C - 25°C = 75°C$

$T_f = 100°C$

$T_i = 25°C$

$q = (435\,g)(4.18\,J/g°C)(75°C)$

$\boxed{q = 136372.5\,J}$ (B)

$\hookrightarrow 1.4 \times 10^5\,J$

13. $q = m H_f$

$q = ?$

$m = 5.00\ g$

$H_f = 334\ J/g$

$q = (5.00\ g)(334\ J/g)$

$q = 1670\ J$ (B)

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14. $q = ?$

$m = 50.0\ g$

$C = 4.18\ J/g°C$

$\Delta T = 60°C - 70°C = -10°C$

$T_f = 60°C$

$T_i = 70°C$

$q = (50g)(4.18\ J/g°C)(-10°C)$

$q = -2090$

$\hookrightarrow -2100\ J$ (C)

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15. $q = m H_v$

$q = ?$

$m = 10.00\ g$

$H_v = 2260\ J/g$

$q = (10.00\ g)(2260\ J/g)$

$q = 22600\ J$

$\downarrow$

$22.6\ KJ$ (D)