CHAPTER 3

(2) states (4) containers (6) gaseous or gas (8) stronger

(10) easily compressed, whereas liquids & solids are not
shape is easily changed, like liquids but unlike solids

(12) chemical (14) chemical, though it sounds like a chemical change.

(16) H_2O molecules split up. The H and O ~~atoms~~ atoms re-combine to make H_2 molecules and O_2 molecules. This is a chemical change because the products (O_2 & H_2) have different chemical properties than water. For instance, water can't burn. The new substances (O_2 & H_2) can not be turned back into H_2O without using a chemical reaction.

(18)
(a) chem change
(b) arguably chem or phys change. I would go with chemical.
(c) chem change
(d) chem change
(e) chem change
(f) chem change
(g) phys change
(h) phys change
(i) chem change
(j) phys change
(k) chem change

(20) element (22) elements (24) different

(26) skip this, bad question

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(28) heterogeneous

(30) (a) could go either way
(b) pure substance
(c) mixture due to stuff dissolved or floating in the water
(d) pure substance

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(32) (a) homogeneous (b) heterogeneous (c) heterogeneous (d) homogeneous
bad question(?) → (e) heterogeneous because of inks

(34) filtration, maybe evaporation

(36) If two or more substances in a solution have different boiling points, they can be separated by distillation. Water boils at 100°C, and salt boils at a <u>much</u> higher temperature. Thus, the water leaves the mixture & the salt stays behind, physical change: you could recover the original mixture by dumping the salt & water together again.

(38) one calorie (40) the molecules move so fast that their motion energy is greater than the energy holding ("attractive forces") the molecules together

(42) $Q = m \times C \times \Delta T$ → the **temperature** change

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(44) $Q = 526\ J$
$m = 7.40\ g$
$\Delta T = 17°C$ ← 2 sig figs

well, 1.052 kJ = 1,052 J.

$\frac{1{,}052\ J}{526\ J} = 2$. So, you're using twice as much energy. So, you could heat up twice as much water: ~~[illegible]~~ 14.8 g **= 15 g**

Easier way!

$Q = mC\Delta T$

$1052\ J = m \left(\frac{4.18\ J}{g\,°C}\right)(17°C)$

$\frac{(1052)}{(4.18)(17)} = m =$ ~~14.8 J~~ $=$ **15 g** → 2 sig figs

(46) 1 cal = 4.184 J thus 1 kcal = 4.18 kJ

(a) $462.4\ kJ \times \frac{1\ kcal}{4.184\ kJ} =$ **110.5 kcal**

(b) $18.28\ kJ \times \frac{1\ kcal}{4.184\ kJ} =$ **4.369 kcal**

(c) $1.014\ kJ \times \frac{1\ kcal}{4.184\ kJ} =$ **0.2424 kcal**

(d) $190.5\ kJ \times \frac{1\ kcal}{4.184\ kJ} =$ **45.53 kcal**

(48) (a) $12.30\ kcal \times \frac{10^3\ cal}{1\ kcal} = 1.230 \times 10^4\ cal$

(b) $290.4\ kcal \times \frac{10^3\ cal}{1\ kcal} = 2.904 \times 10^5\ cal$

(c) $940{,}000\ kcal \times \frac{10^3\ cal}{1\ kcal} = 9.4 \times 10^8\ cal$

(d) $4201\ kcal \times \frac{10^3\ cal}{1\ kcal} = 4.201 \times 10^6\ cal$

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(50) (a)

Step one $75.25\ kJ = \_\_\_\ cal$

Step two $75.25\ kJ \times \frac{J}{kJ} \times \frac{cal}{J} = \_\_\_\ cal$

Step three $75.25\ kJ \times \frac{10^3\ J}{1\ kJ} \times \frac{1\ cal}{4.184\ J} =$ ~~18000~~ 17985.18164 cal → $\underline{17990}$

$1.799 \times 10^4\ cal$

(b) $133200\ cal = \_\_\_\ kJ$

$133200\ cal \times \frac{4.184\ J}{1\ cal} \times \frac{1\ kJ}{10^3\ J} =$ **557.3 kJ**

(c) $232.4\ kcal \times \frac{10^3\ cal}{1\ kcal} \times \frac{4.184\ J}{1\ cal} =$ **$9.724 \times 10^5\ J$**

(d) $645{,}200\ J \times \frac{1\ cal}{4.184\ J} \times \frac{1\ kcal}{10^3\ cal} =$ **154.2 kcal**

(52) $Q = mC\Delta T$

$52.5\ kJ = 52{,}500\ J \quad 1.02\ kg \times \frac{10^3\ g}{1\ kg} = 1020\ g$

$52{,}500\ J = (1020\ g)(C)(11.2°C)$

$$\frac{52{,}500\ J}{(1020\ g)(11.2°C)} = C$$

$$\frac{52{,}500\ J}{11{,}424\ g°C} = C$$

$4.60\ \frac{J}{g°C} = C$

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(54) $Q = 1.25\ kJ = 1{,}250\ J$

$\Delta T = T_{final} - T_{initial} = 15.2°C - 12.0°C = 3.2°C$

C of pure silver $= 0.24 \frac{J}{g°C}$ (from table 3.2 on p. 72)

$m = ?$

$$Q = mC\Delta T \implies m = \frac{Q}{C\Delta T} = \frac{1250\ J}{(0.24 \frac{J}{g°C})(3.2°C)}$$

$$= 1627.6\ g = 1600\ g = 1.6 \times 10^3\ g$$

(56) $Q = mC\Delta T$ Find ΔT $\implies \Delta T = \frac{Q}{mC}$

$Q = 4.52 \times 10^3\ J$

$m = 742.1\ g$

$C = 0.45 \frac{J}{g°C}$ (from table 3.2 on p. 72)

$$\Delta T = \frac{Q}{mC} = \frac{4.52 \times 10^3\ J}{(742.1\ g)(0.45 \frac{J}{g°C})} = 13.535°C = 14°C$$

temp will go up by 14°C

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(58) $C = 0.13 \frac{J}{g°C}$ convert to $\frac{cal}{g°C}$

SKIP the second part of this problem!!

$$\frac{0.13\ J}{g°C} \times \frac{1\ cal}{4.18\ J} = 0.031 \frac{cal}{g°C}$$

...please SKIP second part of problem...

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(60)

25.0 g gold

$m = 25.0\ g$

From p. 72 table 3.2: $C = 0.13\ \frac{J}{g°C}$

$\Delta T = 20.°C$

$Q = mC\Delta T$

$Q = 25.0g \times 0.13 \frac{J}{g°C} \times 20°C$

$Q = 65\ J$

25.0 g mercury

$m = 25.0g$

$C = 0.14\ \frac{J}{g°C}$

$\Delta T = 20.°C$

$Q = mC\Delta T$

$Q = 25.0g \times 0.14 \frac{J}{g°C} \times 20°C$

$Q = 70\ J$

25.0 g carbon

$m = 25.0g$

$C = 0.71\ \frac{J}{g°C}$

$\Delta T = 20.°C$

$Q = mC\Delta T$

$Q = (25.0g \times 0.71 \frac{J}{g°C} \times 20°C)$

$Q = 360\ J$

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(62)

$m = 35.2g$

$Q = 1251\ J$

$\Delta T = 25.0°C$

$C = ?$

$Q = mC\Delta T$

$\frac{Q}{m\Delta T} = C$

$\frac{1251 J}{(35.2g) \times (25.0°C)} = C$

$1.42\ \frac{J}{g°C} = C$